The Periodic Table and Periodicity

Matthew Williams • Chemistry • May 15, 2026

The Periodic Table and Periodicity

The periodic table organises every known element in a way that reveals patterns. Once you understand why elements are arranged as they are, you can predict the behaviour of unfamiliar elements without memorising individual facts.

Historical Development

Several scientists contributed to the periodic table before its modern form was established.

Johann Döbereiner (1817–1829) observed that certain elements could be grouped into sets of three, called triads, where the middle element had properties and an atomic mass approximately halfway between the other two. For example, lithium (7), sodium (23), and potassium (39) form a triad: the average of 7 and 39 is 23, matching sodium. This was an early indication that element properties followed a pattern, but the system only worked for a handful of groups.

Dmitri Mendeleev (1869) arranged elements in order of increasing atomic mass, placing elements with similar properties in the same vertical column. He left deliberate gaps where undiscovered elements should fit and predicted their properties. When those elements were later found, his predictions proved largely correct, which established the table's credibility.

Henry Moseley (1914) demonstrated that elements should be ordered by increasing atomic number (proton number) rather than atomic mass. This resolved anomalies in Mendeleev's arrangement and produced the modern periodic law: the physical and chemical properties of elements are periodic functions of their atomic numbers.

Structure of the Periodic Table

The modern periodic table organises elements into:

- **Periods** — horizontal rows. All elements in the same period have the same number of occupied electron shells.
- **Groups** — vertical columns. All elements in the same group have the same number of valence electrons, which gives them similar chemical properties.

The group number tells you the number of valence electrons; the period number tells you how many electron shells are occupied.

📋 **Example**

Magnesium has the configuration 2,8,2:

- 2 valence electrons 'Group II
- 3 occupied shells 'Period 3

Chlorine has the configuration 2,8,7:

- 7 valence electrons 'Group VII
- 3 occupied shells 'Period 3

📝 **Exam Tip**

Do not confuse group and period. Group = number of valence electrons = vertical column. Period = number of occupied shells = horizontal row.

Trends in Group II (Alkaline Earth Metals)

Group II elements — beryllium, magnesium, calcium, strontium, barium — all have 2 valence electrons and form 2+ ions.

Ease of Ionisation

Ionisation (removal of valence electrons) becomes easier going down Group II. As atomic radius increases down the group, the outer electrons are further from the nucleus and more shielded from its attraction by inner shells. Less energy is therefore needed to remove them.

Reactivity with Oxygen

All Group II metals burn in oxygen to form metal oxides. Reactivity increases down the group:

$$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$$

$$2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$$

Reactivity with Water

Magnesium reacts very slowly with cold water but burns vigorously in steam:

$$Mg + H_2O(g) \rightarrow MgO + H_2$$

Calcium reacts steadily with cold water, producing calcium hydroxide and hydrogen gas:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

Moving down the group, reactivity with cold water increases progressively.

Reactivity with Dilute Hydrochloric Acid

Magnesium and calcium both react with dilute hydrochloric acid to produce a salt and hydrogen gas. The reaction rate increases down the group:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

$$Ca + 2HCl \rightarrow CaCl_2 + H_2$$

> 🧠 **Remember**
> The overall trend in Group II: reactivity increases going down because valence electrons are held less tightly as atomic radius increases and shielding from inner shells increases.

Trends in Group VII (Halogens)

Group VII — fluorine, chlorine, bromine, iodine, astatine — all have 7 valence electrons. They gain one electron to form −1 ions and are the most reactive non-metals.

Physical States at Room Temperature

Element	Formula	State	Colour
Fluorine	F_2	Gas	Pale yellow
Chlorine	Cl_2	Gas	Yellow-green

Bromine	Br,	Liquid	Red-brown
Iodine	I,	Solid	Grey-black

Moving down the group, the molecules become heavier and intermolecular forces between them strengthen, raising melting and boiling points.

Oxidising Power

Halogens are oxidising agents — they gain electrons from other substances. Oxidising power **decreases** going down Group VII. Fluorine is the strongest oxidising agent of all known elements; iodine is considerably weaker.

As atomic radius increases down the group, the outer shell is further from the nucleus and more shielded by inner shells. The ability to attract an incoming electron weakens.

Displacement Reactions

A more reactive halogen displaces a less reactive one from a solution of its salt, which directly demonstrates the relative oxidising powers:

$$\mathrm{Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)}$$

Chlorine displaces bromine because it is a stronger oxidising agent. Bromine can displace iodine for the same reason, but bromine cannot displace chlorine.

> 📝 **Exam Tip**
> In Group VII, reactivity decreases down the group — the opposite direction to Group I and Group II. This is because halogens react by **gaining** electrons, not losing them, and the ability to attract electrons weakens as atomic size increases.

Trends Across Period 3

Period 3 contains eight elements: Na, Mg, Al, Si, P, S, Cl, and Ar. Moving from left to right, the elements change from metals through a metalloid to non-metals:

Element	Symbol	Z	Type	Valence electrons
Sodium	Na	11	Metal	1
Magnesium	Mg	12	Metal	2

Aluminium	Al	13	Metal	3
Silicon	Si	14	Metalloid	4
Phosphorus	P	15	Non-metal	5
Sulfur	S	16	Non-metal	6
Chlorine	Cl	17	Non-metal	7
Argon	Ar	18	Noble gas	8

As nuclear charge increases across the period (more protons, same number of shells), atomic radius decreases and ionisation energy generally rises, making it progressively harder to remove electrons. This shift produces the transition from metallic to non-metallic character.

Oxides Across Period 3

The type of oxide formed by each element also changes systematically:

Element	Oxide	Character
Na	NaO̦	Basic
Mg	MgO	Basic
Al	AlO̦ƒ	Amphoteric (reacts with both acids and bases)
Si	SiO,	Weakly acidic
P	PO̦•€	Acidic
S	SOƒ	Acidic
Cl	ClO̦‡	Acidic

Basic oxides react with acids; acidic oxides react with bases; amphoteric oxides (aluminium oxide) react with both. This trend from basic to amphoteric to acidic directly mirrors the transition from metal to non-metal across the period.

Predicting Properties from Position

Once the group and period of an element are known, many of its properties follow:

- **Group number** ’valence electrons ’likely ion charge and bonding type
- **Period number** ’number of electron shells ’relative atomic size
- **Left of table (low group number)** ’metal, loses electrons, forms basic oxide, positive ion

- **Right of table (high group number)** 'non-metal, gains or shares electrons, forms acidic oxide, negative ion

📋 **Example**

An unknown element X is in Group II, Period 4. Expected properties:

- 2 valence electrons; forms X^{2+} ion
- Metal with a basic oxide
- More reactive than calcium (larger atomic radius, valence electrons lost more easily)
- Reacts with water to produce hydrogen gas and a metal hydroxide
- Reacts with dilute acids to produce a salt and hydrogen gas